

Chapter-10

Lesson-1

I – Characteristics of Gases

- A **gas** is a phase of matter that has no definite shape or volume.
- A gas expands (or is compressed) to completely fill the volume of its container. Thus, a gas takes the volume of its container.
- A gas takes the shape of its container.
- A gas is highly compressible.
- A gas will **homogeneously** mix with any other gas.
- A gas that is ordinarily a liquid or solid at ordinary conditions is called a **vapor**.
- The particles of a gas are extremely far apart in relation to the size of the particles themselves and thus a gas has a very low density.
- A gas's density is measured in units of grams per liter (g/L).
- Air is a mixture of gases (*78% N₂, 21% O₂, 0.9% Ar, and trace amounts of other gases including CO₂ and H₂O*).

A) Gas Pressure:

- A gas exerts an outward pressure when its particles collide with the walls of its container.
- Pressure is defined as a force exerted on an area.

$$P = \frac{F}{A}$$

- **Standard atmospheric pressure**, which is the average normal air pressure at sea level, is ...
 - 1.000 atm
 - 29.92 in. Hg
 - 760.0 mmHg
 - 760.0 torr
 - 101.3 kPa
 - 14.7 lb/in² (14.7 psi)

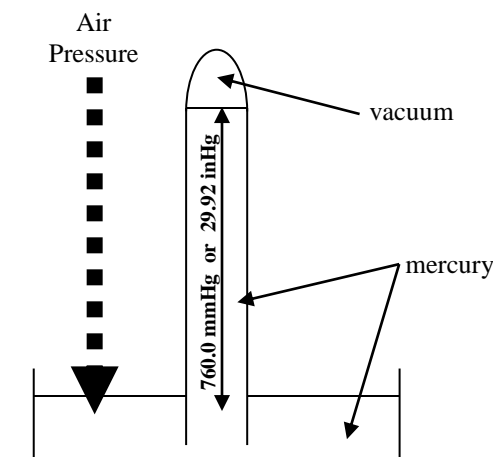
Q1: If air pressure is 30.45 in. Hg, then what is the pressure in torr?

A1: $30.45 \text{ in. Hg} \times \frac{760.0 \text{ torr}}{29.92 \text{ in. Hg}} = 773.5 \text{ torr}$

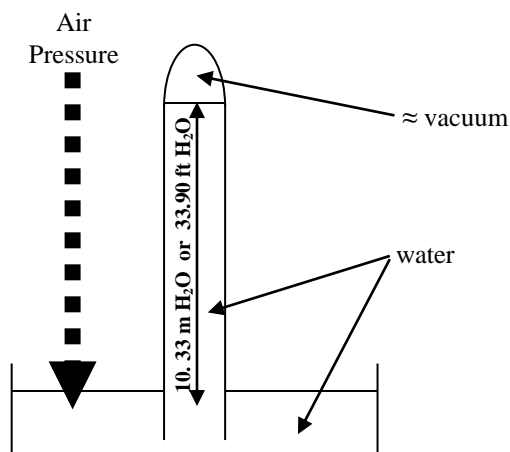
B) Measuring Gas Pressure:

a) The Mercury Barometer and the Water Barometer:

A barometer is used to measure the air pressure.



The height of the mercury column is a measure of the air pressure [normally 29.92 in. at sea level]



The height of a column of water at normal air pressure would be 33.90 feet [water is 13.6 times less dense than mercury, so the water column is 13.6 times higher]

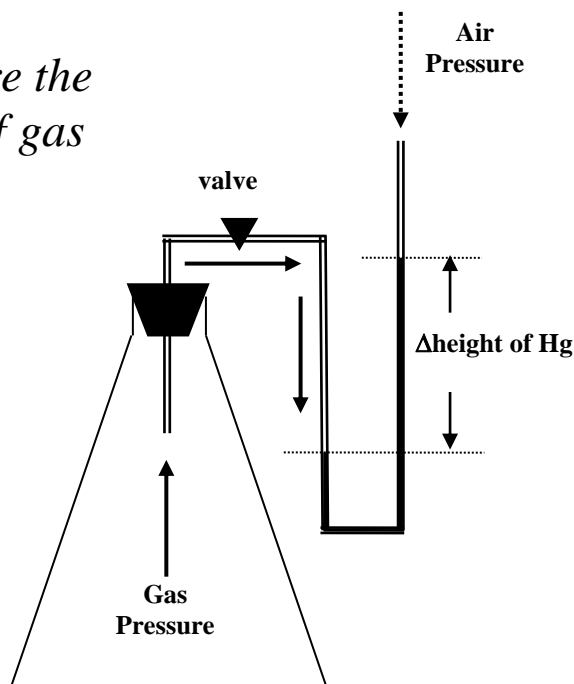
In 1643, Evangelista Torricelli invented the first mercury barometer and created the first known vacuum in science.

b) The Mercury Manometer:

A manometer is used to measure the pressure of a trapped sample of gas

Q2: If the air pressure is 773.5 torr and the height of the mercury is 9.3 mm higher in the open arm of the manometer at the right, then what is the pressure of the gas in the flask?

A2: $773.5 \text{ torr} + 9.3 = 782.8 \text{ torr}$



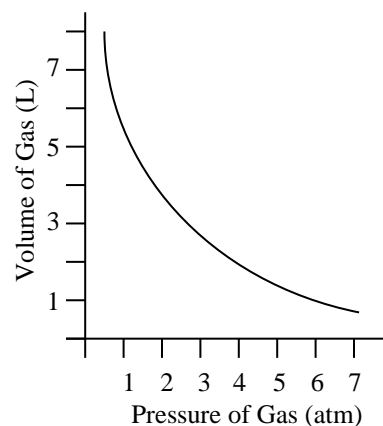
II – The Gas Laws

A) Boyle's Law:

- *The volume of a fixed amount of gas maintained at a constant temperature is inversely proportional to the pressure.*

- $V \propto \frac{1}{P}$ or $V = \frac{k}{P}$

- $P_1V_1 = P_2V_2$



$$V = k \div P$$

at constant n & T

Q3: What will be the volume of 3.00 L of helium gas at standard pressure if the pressure is lowered to 500.0 torr?

A3: $P_1V_1 = P_2V_2$

$$(760.0)(3.00) = (500.0)(X)$$

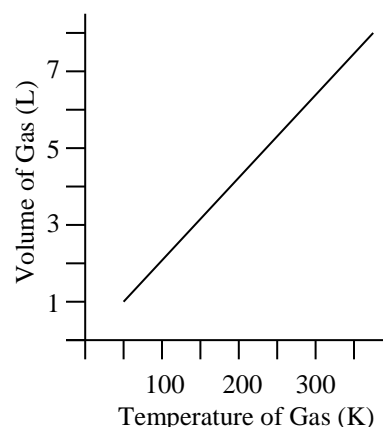
$$X = 4.56 \text{ L}$$

B) Charles's Law:

- *The volume of a fixed amount of gas maintained at a constant pressure is directly proportional to the absolute (Kelvin) temperature.*

- $V \propto T$ or $V = k \times T$

- $\frac{V_1}{T_1} = \frac{V_2}{T_2}$



$$V = kT$$

at constant n & P

Q4: What will be the volume of 3.00 L of helium gas at 20.0°C if the temperature is raised to 93.3°C ?

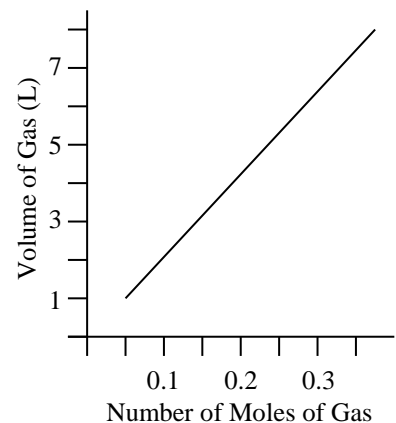
A4: $V_1/T_1 = V_2/T_2$

$$(3.00)/(293.0) = (X)/(366.3)$$

$$X = 3.75 \text{ L}$$

C) Avogadro's Law:

- In 1807, Avogadro stated his hypothesis: *Equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.*
- Avogadro's hypothesis lead to Avogadro's law: ***The volume of a gas maintained at constant temperature and pressure is directly proportional to the number of moles of gas.***



$$V = kn$$

at constant T & P

- $V \propto n$ or $V = k \times n$

- $\frac{V_1}{n_1} = \frac{V_2}{n_2}$

Q5: What will be the volume of 3.00 L of 0.125 moles of helium gas if 0.025 moles of gas escapes?

A5: $V_1/n_1 = V_2/n_2$

$$(3.00)/(0.125) = (X)/(0.100)$$

$$X = 2.40 \text{ L}$$

D) The Combined Gas Law:

- $\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$

Q6: If a piston with 0.0250 mol of gas at 20.0°C and 2.67 atm expands from 225 mL to 1025 mL as the temperature increases to 102.0°C and the amount of gas is increased to 0.0340 mol, then what would be the new pressure?

A6: $P_1 V_1 / n_1 T_1 = P_2 V_2 / n_2 T_2$

$$(2.67)(225) / (0.0250)(293) = (X)(1025) / (0.0340)(375)$$

$$X = 1.02 \text{ atm}$$

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Lesson-2

III – The Ideal-Gas Equation

- An ***ideal gas*** is one in which the individual gas molecules have no volume and have no attractions for each other. Thus, an ideal gas can be compressed into a container with no volume and an ideal gas can never be condensed.
- An ideal gas will perfectly conform to the following relationship called the ***ideal-gas equation***.

$$PV = nRT$$

“pivnert”!

- where P is pressure of gas in atm
 - where V is volume of gas in liters
 - where n is amount of gas in moles
 - where R is the constant 0.08206 L·atm/mol·K
 - [R = 0.08206 L·atm/mol·K = 8.314 L·kPa/mol·K = 8.314 J/mol·K]
- Recall that PV is an energy relationship. Thus “L·kPa” = J.***
- where T is temperature in kelvin

Q1: What is the mass of 3.00 L of He at 20.0°C and 1.05 atm?

A1: $PV = nRT$

$$(1.05)(3.00) = (x)(0.08206)(293.0)$$

$$x = 0.131 \text{ mol} \times \frac{4.00 \text{ g}}{1 \text{ mol}} = 0.524 \text{ g He}$$

Q2: What is the pressure of 454 g of O₂ in a 100.0 L container at 20.0°C?

A2: $454 \text{ g O}_2 \times \frac{1 \text{ mole}}{32.0 \text{ g}} = 14.2 \text{ mol O}_2$

$$PV = nRT$$

$$(x)(100.0) = (14.2)(0.08206)(293.0) \rightarrow x = 3.41 \text{ atm}$$

- *The conditions 273 K and 1.00 atm are referred to as the standard temperature and pressure (STP).*

Q3: What is the volume of 1.00 mole of an ideal gas at STP?

A3: $PV = nRT$

$$(1.00)(x) = (1.00)(0.08206)(273)$$

$$x = 22.4 \text{ L}$$

- *Real gases do not match the ideal gas behavior exactly but under normal condition the differences are usually negligible.*

<i>Molar Volumes of Various Gases at STP</i>						
<i>Ideal gas</i>	Cl_2	CO_2	NH_3	N_2	He	H_2
22.41 L	22.06 L	22.31 L	22.40 L	22.40 L	22.41 L	22.42 L

IV – Further Applications of the Ideal-Gas Equation

- The ideal gas equation can be rewritten as ...

$$\frac{n}{V} = \frac{P}{RT}$$

and both sides can be multiplied by the molar mass (\mathcal{M}).

$$\frac{n\mathcal{M}}{V} = \frac{P\mathcal{M}}{RT}$$

Since $n\mathcal{M}$ equals to grams (g), then ...

$$\frac{g}{V} = \frac{P\mathcal{M}}{RT}$$

and since g/V equals to density (d), then ...

$$d = \frac{P\mathcal{M}}{RT}$$

and it can be rewritten as ...

$$dRT = \mathcal{M}P$$

“dirty imp”!

Q4: What is the density of oxygen gas at 735.0 mmHg and 27.0°C?

A4: $dRT = MP$

$$(x)(0.08206)(300.0) = (32.0)(735.0/760.0)$$

$$x = 1.26 \text{ g/L}$$

Q5: What is the molar mass of an elemental gas whose density in Denver (0.800 atm and 21.0°C) is 1.26 g/L? What is the gas?

A5: $dRT = MP$

$$(1.26)(0.08206)(294.0) = (x)(0.800)$$

$$x = 38.0 \text{ g/mol} \rightarrow \text{It's Fluorine gas!}$$

- *The ideal-gas law can be used to solve stoichiometry problems involving a gaseous reactant or product.*

Q6: What mass of zinc is required to react with excess acid to form 16.0 L of hydrogen gas at STP?

A6: $\text{Zn}_{(s)} + 2 \text{H}^+_{(aq)} \rightarrow \text{Zn}^{2+}_{(aq)} + \text{H}_{2(g)}$

$$16.0 \text{ L H}_2 \times \frac{1 \text{ mole}}{22.4 \text{ L}} \times \frac{1 \text{ mol Zn}}{1 \text{ mol H}_2} \times \frac{65.39 \text{ g}}{1 \text{ mol}} = 46.7 \text{ g}$$

Q7: What volume of oxygen gas at STP would exactly react with 10.8 g of white phosphorus?

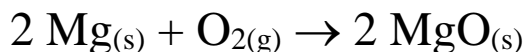
A7: $\text{P}_{4(s)} + 5 \text{O}_{2(g)} \rightarrow \text{P}_4\text{O}_{10(g)}$

$$10.8 \text{ g P}_4 \times \frac{1 \text{ mole}}{124 \text{ g}} \times \frac{5 \text{ mol O}_2}{1 \text{ mol P}_4} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 9.75 \text{ L}$$

Q8: What mass of magnesium can be completely burned by 1.00 L O₂ at 1000.°C and 0.900 atm?

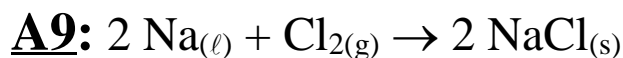
A8: $(0.900)(1.00) = (x)(0.08206)(1273)$

$$x = 0.00862 \text{ mol O}_2$$



$$0.00862 \text{ mol O}_2 \times \frac{2 \text{ mol Mg}}{1 \text{ mol O}_2} \times \frac{24.3 \text{ g}}{1 \text{ mol}} = 0.419 \text{ g Mg}$$

Q9: How many liters of chlorine gas at 1.10 atm and 200.°C would be needed to exactly react with 5.00 g of molten sodium?



$$5.00 \text{ g Na} \times \frac{1 \text{ mol}}{23.0 \text{ g}} \times \frac{1 \text{ mol Cl}_2}{2 \text{ mol Na}} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 2.43 \text{ L} \leftarrow \text{This is at STP!}$$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$\frac{(1.00)(2.43)}{(273)} = \frac{(1.10)(x)}{(473)}$$

$$x = 3.83 \text{ L Cl}_2$$

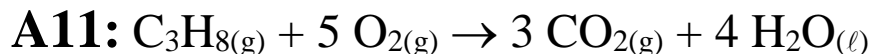
- *When two different gases are involved in a reaction, the coefficients in the balanced equation can be interpreted as volumes (instead of moles) assuming that temperature and pressure is constant.*
- *This is known as Gay-Lussac's Law of Combining Volumes.*

Q10: What volume of nitrogen gas at STP would be needed to react with 6.72 L of hydrogen gas to produce ammonia?



$$6.72 \text{ L H}_2 \times \frac{1 \text{ L N}_2}{3 \text{ L H}_2} = 2.24 \text{ L N}_2$$

Q11: What volume of oxygen gas at 1.00 atm and 20.0°C would be needed to **completely** burn an entire 50.0 L tank of propane gas under 5.00 atm pressure?



$$P_1 V_1 = P_2 V_2$$

$$(5.00 \text{ atm})(50.0 \text{ L}) = (1.00 \text{ atm})(x)$$

$$x = 250. \text{ L C}_3\text{H}_8 \text{ [this is the volume of the propane at 1 atm.]}$$

$$250. \text{ L C}_3\text{H}_8 \times \frac{5 \text{ L O}_2}{1 \text{ L C}_3\text{H}_8} = 1250 \text{ L O}_2 \text{ at 1.00 atm}$$

Chapter-10

Lesson-3

V – Gas Mixtures and Partial Pressures

- The total pressure of a gas mixture in a rigid container is the sum of the partial pressures of the component gases of the mixture at constant temperature and constant volume.

$$P_{\text{tot}} = P_1 + P_2 + \dots + P_n$$

- The total volume of a gas mixture is the sum of the partial volumes of the component gases of the mixture at constant pressure and constant temperature.

$$V_{\text{tot}} = V_1 + V_2 + \dots + V_n$$

- The total moles of gas in a gas mixture are the sum of the moles of the component gases of the mixture.

$$n_{\text{tot}} = n_1 + n_2 + \dots + n_n$$

- In a gas mixture, the ratio of the pressure of Gas-1 (or 2 or 3) alone to the total pressure of the mixture is equal to the ratio of the volume of Gas-1 (or 2 or 3) alone to the total volume of the mixture.

$$\frac{P_{\text{tot}}}{V_{\text{tot}}} = \frac{P_1}{V_1} \quad \text{or} \quad \frac{P_1}{P_{\text{tot}}} = \frac{V_1}{V_{\text{tot}}}$$

- In a gas mixture, the ratio of the pressure of Gas-1 (or 2 or 3) alone to the total pressure of the mixture is equal to the ratio of the moles of Gas-1 (or 2 or 3) alone to the total moles of the mixture.

$$\frac{P_{\text{tot}}}{n_{\text{tot}}} = \frac{P_1}{n_1} \quad \text{or} \quad \frac{P_1}{P_{\text{tot}}} = \frac{n_1}{n_{\text{tot}}}$$

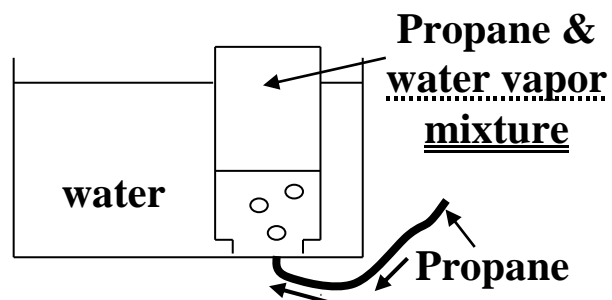
- ***In a gas mixture, the pressure of Gas-1 alone is called the partial pressure of Gas-1.***

- In a gas mixture, the ratio of the moles of Gas-1 to the total moles of gas is called the **mole fraction (X)** of Gas-1.

$$X_1 = \frac{n_1}{n_{\text{tot}}}$$

- The partial pressure of a gas is equal to its mole fraction multiplied by the total pressure.

$$P_1 = X_1 P_{\text{tot}}$$



- Whenever any gas is collected over water, there will always be **some water vapor** in the container with the collected gas. The total pressure of the mixture is equal to the sum of the partial pressures of the gas and water vapor.

$$P_{\text{tot}} = P_{\text{collected gas}} + P_{\text{water}}$$

The water's partial pressure [vapor pressure] is usually found by measuring the water's temperature and then looking up the partial pressure [vapor pressure] on a table. (**Appendix B in textbook**)

Q1: A sample of propane gas is collected over water in a 200.0 mL bottle at 23.0°C and at a total pressure of 712.0 mmHg.

- Calculate the partial pressure of propane gas in the mixture.
- Calculate the moles of propane gas in the mixture.
- Calculate the mole fraction of propane gas in the mixture.

A1: (a) $P_{\text{tot}} = P_{\text{propane}} + P_{\text{water}}$

$$712.0 \text{ mmHg} = x + 21.07 \text{ mmHg} \quad \leftarrow \text{found on Appendix B at } 23.0^\circ\text{C}$$

$$x = 690.1 \text{ mmHg propane}$$

(b) $PV = nRT$

$$(690.1/760.0)(0.2000) = (x)(0.08206)(296.0)$$

$$x = 0.007477 \text{ moles propane}$$

(c) $P_{\text{propane}} = X_{\text{propane}} \cdot P_{\text{tot}}$

$$690.1 \text{ mmHg} = (x)(712.0 \text{ mmHg})$$

$$x = 0.969 \quad \leftarrow \text{ie: mixture} = 96.9\% \text{ propane and } 3.1\% \text{ water vapor}$$

VI - Kinetic Molecular Theory of Gases

A) Postulates of the Kinetic Theory:

- Gases consist of a very large number of particles in continuous, random, straight-line motion.
- The total sum of the individual volumes of each gas particle is negligible relative to the total volume of the gas.
- Attractive and repulsive forces between the gas particles are negligible
- The average energy of translational motion (KE_t) of an individual gas particle is directly related to temperature.

$$KE_t = \frac{1}{2}mv^2 \propto T(K)$$

Where:

- " KE_t " is the average KE of the gas particles in **joules**
- " m " is the mass of the gas particles in **kilograms**
- " v " is the average velocity of the gas particles in **meters/second**
- " T " is the absolute temperature in **kelvin**
- Collisions between the gas particles are elastic - no kinetic energy is converted to heat. Thus the temperature of a gas will remain constant.
- A gas (the aggregate of the particles) exerts pressure as a result of the force associated with wall collisions. These collisions are elastic.

B) Gas Pressure Explained using Kinetic Theory:

- In a container, the pressure (P) exerted by a gas increases as...
 - 1) the number of mole of gas (n) increases because, with more particles, there are more collisions with the walls per unit time.
 - 2) the volume of the container (V) decreases because, the gas particles strike the walls of the container more often due to a shorter distance between the walls.
 - 3) the temperature of the gas (T) increases because, at higher temperatures the gas particles are moving faster and strike the walls of the container... (a) more often and (b) with more force.

C) Effusion of Gases - Graham's Law:

- **Effusion** - the flow of a gas through a small opening in a container.
- **Effusion rate** (*r*) is the distance traveled per unit time.
- **Effusion rate** (*r*) is directly proportional to the average KE of the gas particles.

$$KE_{\text{t of gas A}} = \frac{1}{2} (m_{\text{of gas A}})(v_{\text{of gas A}})^2 \propto T(K)$$

$$KE_{\text{t of gas B}} = \frac{1}{2} (m_{\text{of gas B}})(v_{\text{of gas B}})^2 \propto T(K)$$

If both gases are in the same container (at the same temperature and pressure) then...

$$\frac{1}{2} (m_{\text{of gas A}})(v_{\text{of gas A}})^2 = \frac{1}{2} (m_{\text{of gas B}})(v_{\text{of gas B}})^2$$

$$(m_{\text{of gas A}})(v_{\text{of gas A}})^2 = (m_{\text{of gas B}})(v_{\text{of gas B}})^2$$

$$\frac{(v_{\text{of gas A}})^2}{(v_{\text{of gas B}})^2} = \frac{(m_{\text{of gas B}})}{(m_{\text{of gas A}})}$$

$$\frac{r_{\text{of gas A}}}{r_{\text{of gas B}}} = \left[\frac{\mathcal{M}_{\text{of gas B}}}{\mathcal{M}_{\text{of gas A}}} \right]^{1/2}$$

In order that both gases have the same average KE, then the lighter gas must move faster than the heavier gas and vice versa.

Q2: A certain gas (X) takes 3.546 times longer to effuse than N₂ at the same temperature and pressure. What is the molar mass of the gas (X)?

$$\text{A2: } \frac{r_{\text{of gas X}}}{r_{\text{of gas N}_2}} = \left[\frac{\mathcal{M}_{\text{of N}_2}}{\mathcal{M}_{\text{of X}}} \right]^{1/2}$$

Since the unknown gas is 3.546 times slower than the nitrogen gas, then assign the unknown gas a rate of 1.000 and assign the nitrogen gas the rate of 3.546.

$$\frac{1.000}{3.546} = \left[\frac{28.0}{X} \right]^{1/2}$$

$$X = 352 \text{ g/mol}$$

FYI, the gas is UF₆

- ☞ If two gases are diffusing the same distance, Graham's law can be written, where **t** is time.

$$\frac{\text{t of gas A}}{\text{t of gas B}} = \left[\frac{\mathcal{M} \text{ of gas A}}{\mathcal{M} \text{ of gas B}} \right]^{1/2}$$

Q3: If a bottle of ammonia (NH₃) is uncorked 0.600 m away and it takes 3.00 sec to smell it, then how long should it take to smell an uncorked bottle hydrogen sulfide (H₂S) at the same distance?

A3: $\frac{\text{t of gas NH}_3}{\text{t of gas H}_2\text{S}} = \left[\frac{\mathcal{M} \text{ of NH}_3}{\mathcal{M} \text{ of H}_2\text{S}} \right]^{1/2}$

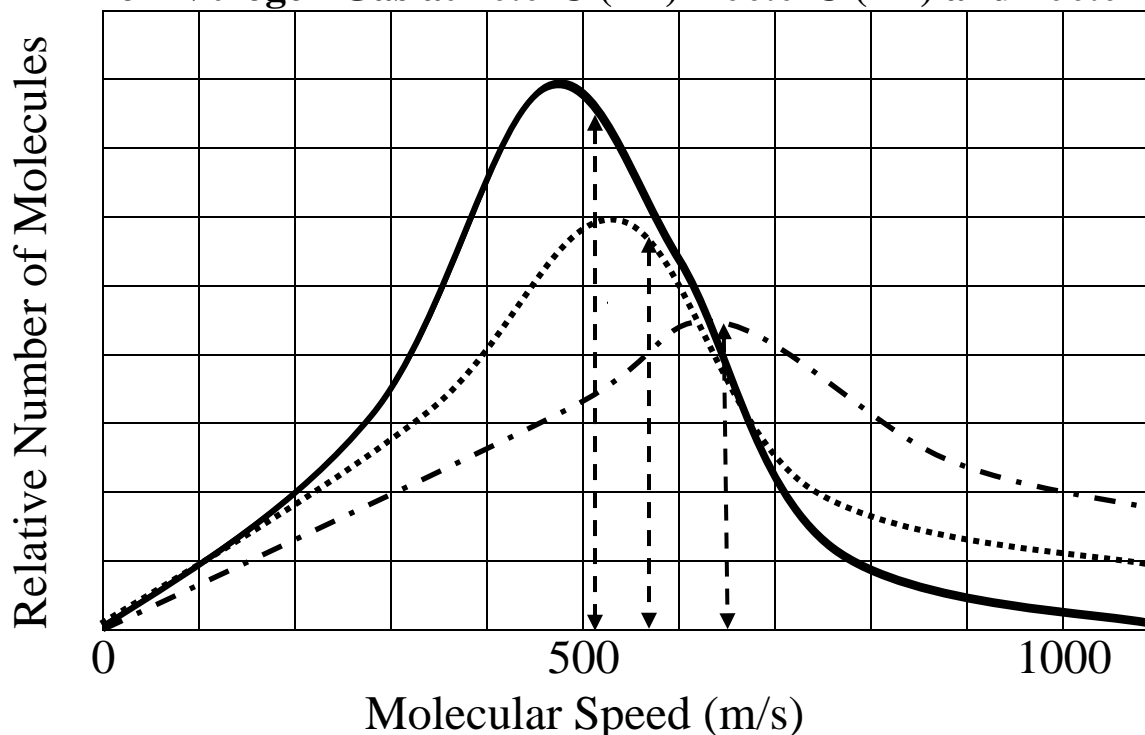
$$\frac{3.00 \text{ s}}{X} = \left[\frac{17.0}{34.1} \right]^{1/2}$$

$$X = 4.25 \text{ s}$$

D) The "Average" Speed of a Gas Particle:

- Recall: $\text{KE}_{\text{t of gas A}} = \frac{1}{2} (\text{m of gas A})(v \text{ of gas A})^2 \propto T(\text{K})$
- Recall: the temperature of a substance is directly proportional to the average translational kinetic energy of its particles.

Maxwell-Boltzmann Distribution of Molecular Speeds for Nitrogen Gas at 20.0°C (—) 100.0°C (.....) and 200.0°C (– ·)



Chapter-10

Lesson-4

VII –The Nonideal Behavior of Real Gases

<u>Ideal Gas Particles</u>	<u>Real Gas Particles</u>
• have no volume	• have a small volume
• have no attractive forces for each other	• have small attractive forces for each other

A) Conditions of High Temp. and Low Pressure:

- The ***difference in behavior*** between a sample of a real gas and a sample of an ideal gas is small at ordinary temperatures and pressures and is very small at high temperatures and low pressures (*usually less than 10 atm*)!

- $PV = nRT$

- $dRT = \mathcal{M}P$

- $\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$

} These laws are able to reasonably predict real gas behavior under ordinary conditions of temperature and pressure!

WHY?

- 1) Under these (low) pressures, the real gas particles are very far apart in relation to their own size (volume). Thus the fact that the particles actually have volume is not very noticeable in relation to the total volume of the sample.
- 2) Under these (low) pressures, the fact that real gases do attract each other is ignored because the forces lessen when the particles are farther apart.
- 3) Under these conditions of (high) temperatures, the particles are moving very fast and the attractive forces between the particles become less able to influence the particles' behavior.

B) Conditions of Low Temp. and High Pressure:

- At low temperatures and high pressures (*usually greater than 10 atm*) the difference in behavior between a sample of a real gas and an ideal gas is noticeably different.

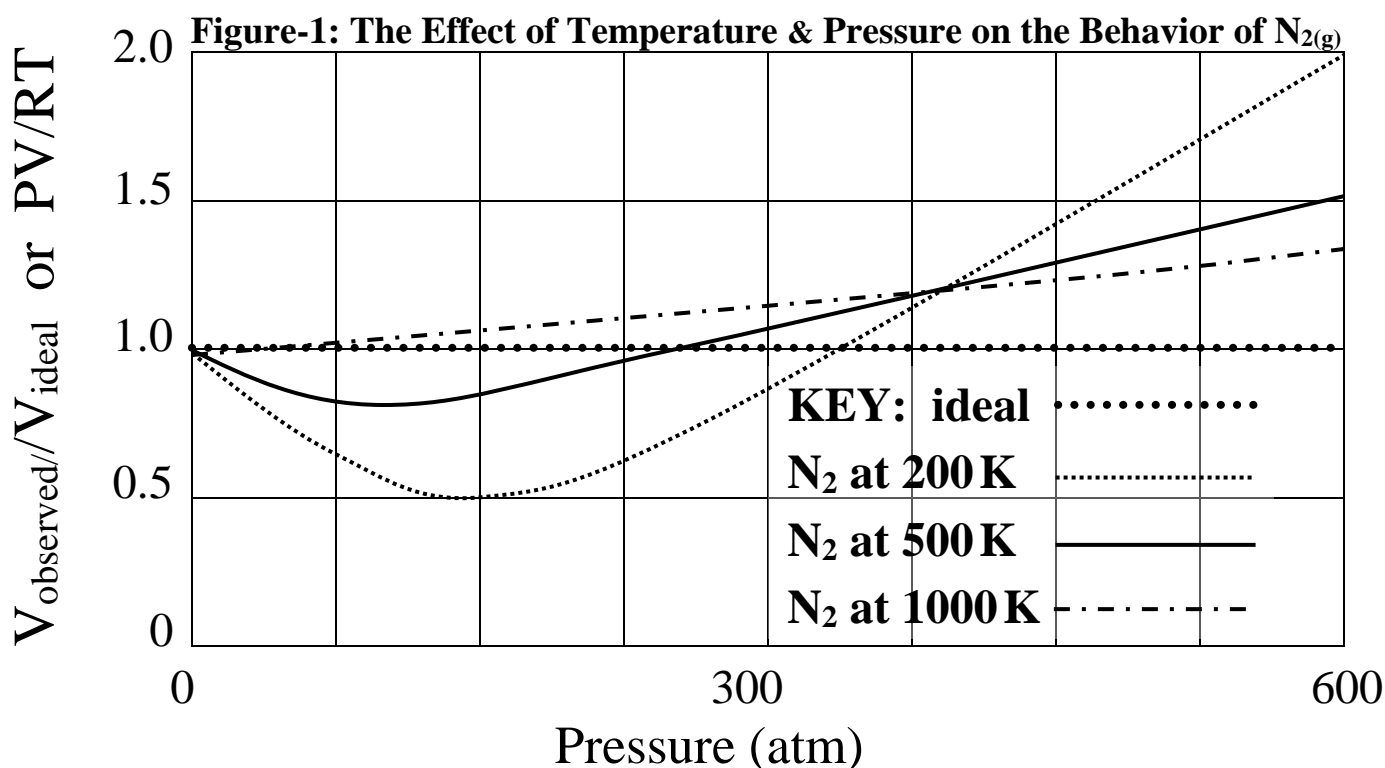
$$\begin{aligned} & - PV = nRT \\ & - dRT = \mathcal{M}P \\ & - \frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \end{aligned}$$



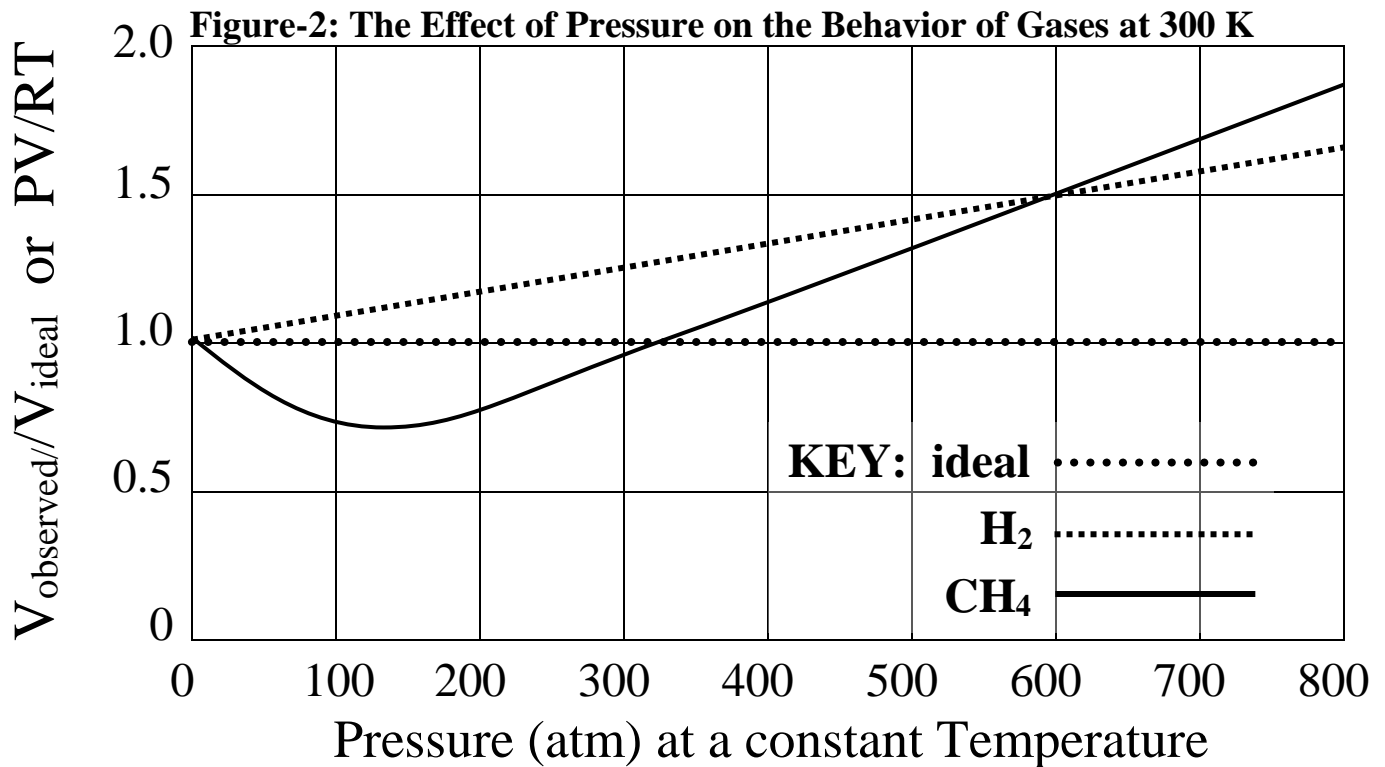
These laws are NOT able to predict real gas behavior under conditions of very low temperature and very high pressure!

WHY?

- Under these (high) pressures, the real gas particles are very close together in relation to their own size (volume). Thus the fact that the particles actually **DO** have volume is very noticeable in relation to the total volume of the sample.
- Under these (high) pressures, the fact that real gases **DO** attract each other becomes important because the attractive forces are stronger when the particles are closer together.
- Under these conditions of (low) temperatures, the particles are moving very slowly and the attractive forces between the particles become more able to influence the particles' behavior.



- Gases with smaller molar masses are generally more ideal their behavior than gases with larger molar masses.



C) The van der Waal's Equation:

$$(P + n^2a/V^2)(V - nb) = nRT$$

In this equation, **a** and **b** are independent of P, V, n, and T. They are specific to the gas, and are called van der Waal's constants.

Gas	"a" ($L^2 \cdot atm/mol^2$)	"b" (L/mol)
He	0.0341	0.0237
H_2	0.244	0.0266
CH_4	2.25	0.0428
N_2	1.39	0.0391

The term " n^2a/V^2 " reflects the attractive forces that DO exist between the gas particles. "P" is the observed gas pressure.

☞ *The term " n^2a/V^2 " adjusts the observed pressure up to ideal pressure!*

The term " nb " reflects the volume that the gas particles DO occupy in the container. "V" is the actual container volume.

"V-nb" is the free (empty) space between the gas particles.

☞ *The term " nb " adjusts the container volume down to ideal volume!*

Q1: Calculate the volume of 1.00 mole of H₂ at 173 K and 10.0 atm using (a) the Ideal Gas Law and (b) the van der Waal's equation.

A1: (a) $PV = nRT$

$$(10.0)(x) = (1.00)(0.0821)(173)$$

$$x = \underline{1.42 \text{ L}}$$

(b) $(P + n^2a/V^2)(V - nb) = nRT$

$$(10.0 + 0.244/x^2)(x - 0.0266) = (1.00)(0.0821)(173)$$

$$x = \underline{1.43 \text{ L}} \quad (101\% \text{ of what the Ideal Gas Law would predict})$$

Q2: Calculate the volume of 1.00 mole of CH₄ at 173 K and 10.0 atm using the van der Waal's equation

A2: $(P + n^2a/V^2)(V - nb) = nRT$

$$(10.0 + 2.25/x^2)(x - 0.0428) = (1.00)(0.0821)(173)$$

$$x = \underline{1.30 \text{ L}} \quad (91.5\% \text{ of what the Ideal Gas Law would predict})$$

Q3: The normal boiling points of CO and SO₂ are -192°C and -10.°C respectively. (a) At 25°C and 1 atm, which gas would you expect to have a molar volume closest to the ideal value?
(b) If you wanted to reduce the deviation from ideal gas behavior, in what direction would you change the temperature and pressure?

A3: (a) carbon monoxide (it's smaller and has weaker IMF's)

(b) Increase the temperature and decrease the pressure.

Q4: Would you expect helium gas to deviate from ideal behavior the most at 100 K and 1 atm, 100 K and 5 atm, or 300 K and 2 atm?

A4: 100 K and 5 atm

Base your answers to the question below on Figure-2.

Q5: (a) Calculate the ideal volume of 1.00 mol of $\text{CH}_4(\text{g})$ at 100 atm and 25°C . (b) Using the graph of the bottom of this lecture's second page, estimate the real volume of $\text{CH}_4(\text{g})$ at 100 atm and 25°C . (c) Explain why the real volume differs from the ideal volume. (d) explain why the V_o/V_i ratio on the graph is greater than 1 at pressures above 350 atm?

A5: (a) $PV = nRT$

$$(100)(x) = (1)(0.0821)(298)$$

$$(x) = 0.245 \text{ L (ideal volume)}$$

(b) The V_o/V_i ratio on the graph is about 0.75.
Real volume is 0.184 L (75% of ideal volume)

(c) At 100 atm, the CH_4 molecules are pushed together and the weak but real attractive forces the CH_4 molecules have for each other dominate and pull in the molecules causing the real total volume to be smaller than ideal behavior would predict.

(d) At 350 atm, the CH_4 molecules are pushed together and the small but real volumes of the individual CH_4 molecules interfere with the gas contracting as it should causing the real total volume to be larger than ideal behavior would predict.